AMSAT Chem 1H Mr. Dehne

Name: _____

Date: _____

Per#: ___

Liquids and Solids Topic#12

WS#1: Properties of Liquids

1. The friction between moving molecules in a liquid is called _____.
2. The stronger the intermolecular forces in a liquid, the _____ the viscosity.
3. Liquids with hydrogen bonds tend to have _____ viscosities.
4. An insect is able to walk on water because of a property known as _____.

T or F. Correct if false.

5. The molecules on a liquid surface are attracted <u>outward</u> and sideways. __ _____
6. Water is an excellent solvent because of the <u>ionic</u> nature of its molecules. __ _____
7. Water is able to rise in plants' roots and stems because of its <u>surface tension</u>. __ _____
8. What effects does an increase in temperature have on viscosity?
9. A drop of mercury on a glass surface beads into a tighter sphere than a drop of water. A drop of alcohol hardly beads at all. Draw each of the drops on the line below.

_____

mercury water alcohol

   a. What can you infer about the intermolecular forces in these three liquids? _____
   _____
   b. Which of these liquids would you predict has the highest viscosity? Why? _____
   _____

10. List three unusual properties of water._____
    _____
11. List the property(ies) of water that account for the following facts:
    a. Water is a liquid at room temperature._____
    b. Oceans moderate the climate on Earth._____
12. How does the structure of water explain its high *bp*, high heat capacity, and high heat of vaporization?
    _____

WS#2: The Nature of Solids

1. Movements among atoms or molecules are limited in both the solid and _____.
2. The smallest repetitive unit in a crystalline structure is called a _____.
3. Unlike crystalline solids, amorphous solids may_____ over a wide range of temperature before melting.
4. In a molecular solid, intramolecular covalent bonds are _____ than intermolecular attractive forces.
5. List 3 of the physical properties of solids that are dependent on both the kind of particles that make up the solids and the strength of the attractive forces between the particles._____
   _____
6. Why are the bonds in sodium chloride stronger than those in a molecular solid such as sugar? _____
   _____
7. What is the principal feature of a metallic bond? _____
   _____
8. Diamond and graphite are both composed of entirely of carbon, yet graphite is soft and a diamond is one of the hardest substances known. Explain the difference between these substances in terms of intermolecular forces.
   _____
9. Which of these two solids would have the higher *mp*: a molecular solid held together by dispersion forces or a molecular solid held together by hydrogen bonds? Why? _____
   _____

a. ionic solid b. metallic solid c. molecular solid d. network-covalent solid

10. ____ high melting point 11. ____ poor conductor 12. ____ malleable 13. ____ brittle

WS#3: Condensed States of Matter

Fill in the following chart comparing solids and liquids.

Physical Property	Liquids	Solids
1. Compressibility		
2. Density		
3. Shape		
4. Rate of Diffusion		

1. Describe the behavior of the molecules in a liquid. Explain this behavior in terms of intermolecular forces.
2. How do intermolecular forces determine molecular arrangement in solids?

If the statement is true, write "true." If false, make true by changing the underlined word.

_____ 7. Intermolecular forces determine <u>metallic</u> properties such as the boiling point of a substance.
_____ 8. Intermolecular forces are forces of attraction between <u>atoms</u>.
_____ 9. Chemical bonds are the forces between <u>molecules.</u>
_____ 10. Covalent bonds result from electrons being shared between atoms in a <u>molecule</u>.
_____ 11. <u>Intermolecular</u> forces result from the electron interactions between neighboring molecules.
_____ 12. When the electron cloud is not distributed symmetrically, a molecule has a <u>polarity</u>.
_____ 13. Chemical bonds include ionic bonds, metallic bonds, and <u>atomic bonds</u>.
_____ 14. Intermolecular forces include dispersion forces, dipole-dipole forces, and <u>helium bond</u> forces.

3. The stronger the intermolecular forces in a liquid, the _____its boiling point.
4. The state of a substance at room temperature depends on_____.
5. Water molecules in an ice cube are held together by _____forces or, more specifically, _____ bonds.

On the line, write the letter of the term that matches each description below. Each choice can be used once, more than once, or not at all. a. London dispersion b. dipole-dipole c. hydrogen bonding

____ 6. The force in between SO_2 molecules
____ 7. The force that accounts for HF being a liquid while H_2 and F_2 are gases
____ 8. The force that depends on temporary, induced dipoles
____ 9. The only intermolecular force that exists in noble gases
____ 10. The attractive force between neighboring dipoles
____ 11. The force that arises because of the large differences in electronegativities in the N-H bonds

Answer the following question as directed.

12. Which of the following statements could be used to explain why water has an unusually high bp? Make a checkmark on the line for all that apply.
    - ____ a. Water molecules have a strong intermolecular force.
    - ____ b. Oxygen is more electronegative than hydrogen.
    - ____ c. Oxygen atoms are smaller than hydrogen atoms.
    - ____ d. There is a large difference in electronegativity between oxygen and hydrogen.
    - ____ e. There are no nitrogen atoms in water.
    - ____ f. Water is an angular molecule.
    - ____ g. Water does not contain metallic bonds.

WS#4: Boiling Point and Melting Point (Intermolecular Forces IMFs))

1.
2.
3.

WS#4: Changes of State

Definitions of abbreviations: Melting point (*mp*), boiling point (*bp*), and freezing point (*fp*)

vaporization / melting / equilibrium vapor pressure / condensation / phase change / volatile / freezing point / sublimation / deposition

1. _____ conversion of a solid into a gas
2. _____ opposite of vaporization
3. _____ temperature at which the solid and liquid forms of a substance exist in equilibrium
4. _____ conversion of a substance from one of the three physical states of matter in equilibrium
5. _____ change of state from liquid to gas
6. _____ pressure exerted by a constant number of gas molecules above a liquid or solid
7. _____ description of a liquid that evaporates easily
8. _____ transformation of a gas directly into a solid
9. _____ phase change from solid to liquid

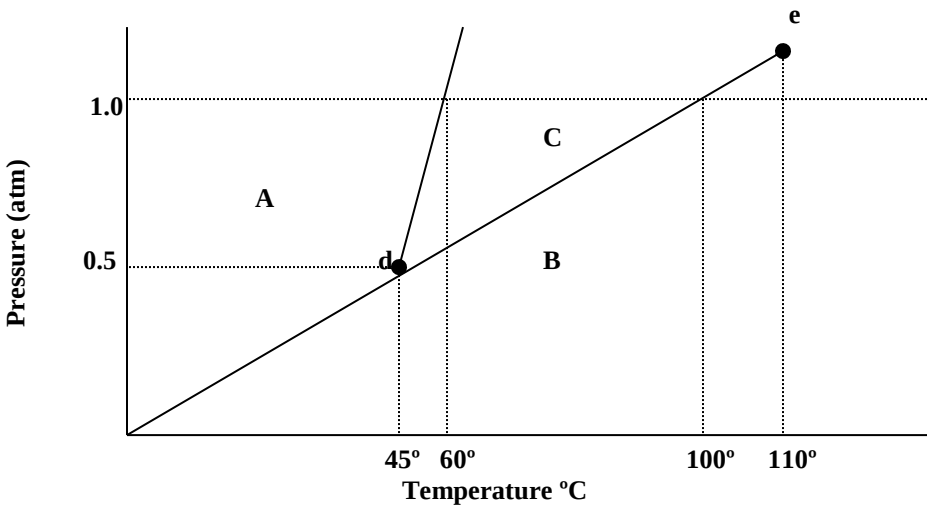
T or F. Correct if false.

10. The KE_{av} of the particles in a liquid depends upon the <u>temperature</u>. __ _____
11. Solids do not flow because the attractive forces between their particles are <u>weaker</u> than those in a liquid or gas. __ _____
12. Dynamic equilibrium is reached when the rate of vaporization <u>exceeds</u> the rate of condensation. __ _____
13. <u>Condensing</u> $H_2O(l)$ is the opposite of melting $H_2O(s)$. __ _____

14. All three states of matter can exist in equilibrium at the <u>triple point.</u> __ _____
15. Why do ice cubes shrink and become rounded if left in the freezer for a long time? _____
_____
16. Six transitions are listed in the diagram below. Use arrows to connect each transition to its beginning state and its ending state. One example is done for you. When completed, each beginning state and ending state will be used twice.

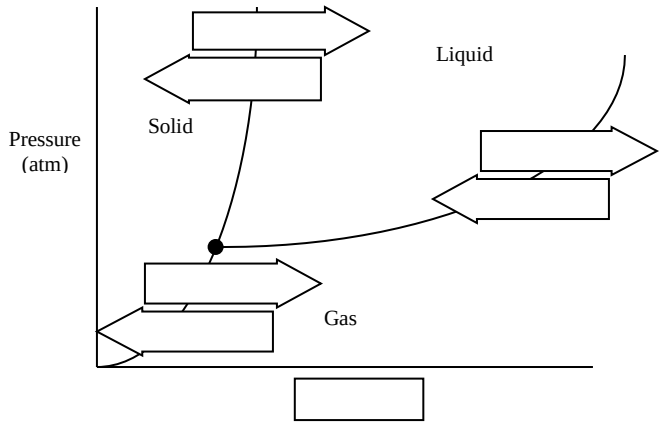
Beginning State	Transition	Ending State
Gas	melting	solid
	freezing	
Liquid	condensation	liquid
	boiling	
Solid	deposition	gas
	sublimation	

WS#5: Phase Diagram

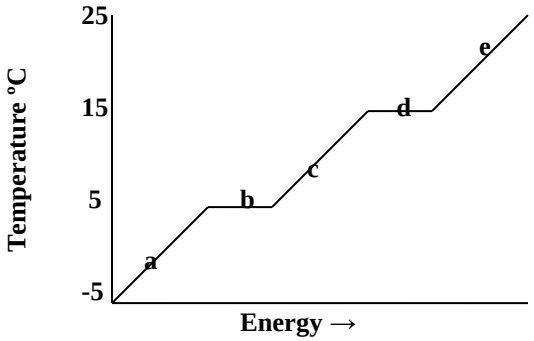


Answer the following questions using the chart above.

1. What section represents the solid phase? _____
2. What section represents the liquid phase? _____
3. What section represents the gas phase? _____
4. What letter represents the triple point? _____
5. What letter represents the critical point? _____
6. What is the substance's normal melting point? _____
7. What is the substance's normal boiling point? _____
8. Above what temperature is it impossible to liquefy this substance regardless of the pressure? _____
9. At what temperature and pressure do all three phases coexist? _____
10. Is the density of the solid greater than or less than the density of the liquid? _____
11. Would an increase in pressure cause this substance to freeze or melt? _____
12. At the triple point of water, ice, liquid water, and water vapor can all exist all at the same time. Compare the KE and *total energy* of the molecules of ice, liquid, and water vapor at the triple point.
13. In the phase diagram below, correctly label the x-axis and the triple point. Write the names of all six phase transitions in the arrows provided.



WS#6: Heating Curve



Answer the following questions using the chart above.

1. What is the *fp* of the substance? _____
2. What is the *bp* of the substance? _____
3. What is the *mp* of the substance? _____
4. What letter represents the range where the solid is being warmed? _____
5. What letter represents the range where the liquid is being warmed? _____
6. What letter represents the range where the vapor is being warmed? _____
7. What letter represents the melting of the solid? _____
8. What letter represents the vaporization of the liquid? _____
9. What letter(s) shows a change in PE? _____
10. What letter(s) shows a change in KE? _____
11. What letter represents condensation? _____
12. What letter represents crystallization? _____

WS#7: Energy Conversions

Convert the following amounts of energy as indicated. Show all work. (4.184J = 1 calorie)

1. 105 J = _____Cal (Ans: 0.0251 Cal)
2. 47,500 cal = _____J (Ans: 1.99×10^5J)
3. A bagel contains _____cal, or 837,000 J. (Ans: 2.00×10^5cal)
4. 0.251 kJ = _____kilocalories (Ans: 0.0600kcal)
5. An apple contains _____cal, or 523 kJ (Ans: 1.25×10^5cal)

Determine the amount of heat (in joules) absorbed /released in each of the following changes. (S_{H2O} = 4.184 J/g-°C, S_{Al} = 0.900 J/g-°C, and S_{Fe} = 0.450 J/g-°C, and S_{Mg} = 1.020)

6. 40.0g of water heated from 10.0ºC to 30.0ºC (Ans: 3.35×10^3J)
7. 25.0g of water cooled from 85.0ºC to 40.0ºC (Ans: -4.71×10^3J)
8. 65.5g of water heated from 32.5ºC to 48.7ºC (Ans: 4.44×10^3J)
9. 30.0g of aluminum heated from 15.0ºC to 35.0ºC (Ans: 530J)
10. 450.0g of iron cooled from 125.0ºC to 45.0ºC (Ans: -1.6×10^4J)
11. 195.4g of magnesium cooled from 120.6ºC to 14.9ºC (Ans: -2.1×10^4J)

Determine the amount of energy needed to transform the following from one phase to another phase.

12. 100.0g of ice melted, with no temperature change (Ans: 33.35kJ)
13. 40.0g of water boiled at 100ºC (Ans: 90.5kJ)

Determine the total amount of heat needed for the following changes.

14. The complete vaporization of 10.0g of ice originally at -10.0ºC. (Ans: 30.4kJ)